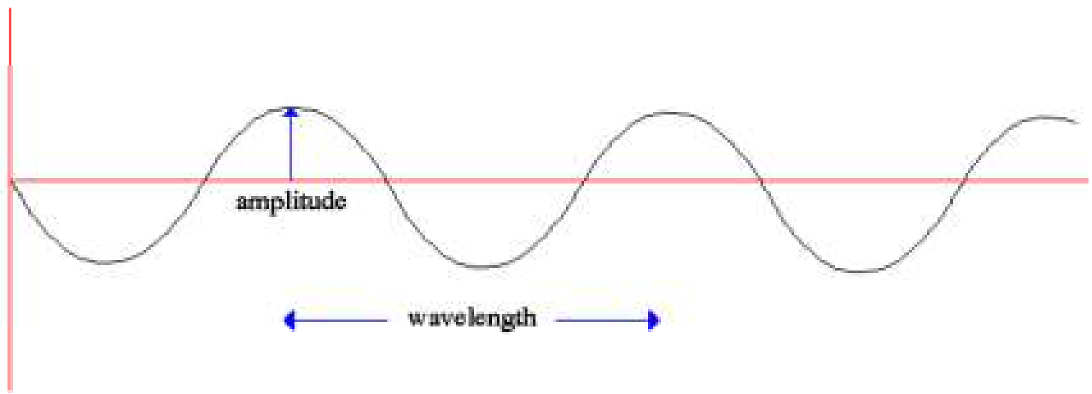
Name ____ AP Chemistry

Electromagnetic Radiation and Atomic Spectra POGIL

Electromagnetic Radiation

Model 1: Characteristics of Waves



The figure above represents part of a wave. The entire wave can be thought of extending infinitely in both directions. One important characteristics of a wave is its wavelength (λ), as marked on the figure above.

1) Based on the figure above, define wavelength.

2) Suppose that the wave depicted above is travelling to the left at a speed of 35 cm/sec and that $\lambda = 2.5$ cm.
   (a) How long would it take for 1 wavelength (or 1 cycle of the wave to travel through the origin of the graph?

   (b) How many wavelengths (or cycles) would travel through the origin during a time interval of 1 second?

   (c) Would your answer change in a and b decrease, increase, or remain the same if $\lambda >$ 2.5 cm? Explain your reasoning.

3) The frequency (f) of a wave is defined as the number of wavelength per second which travel past a given point.
   (a) For a wave travelling at a given speed (c), how does the frequency depend on the wavelength, if at all?

   (b) Provide a mathematical expression showing the relationship between f, c, and λ for a wave. (Hint: consider your answers for 2a and 2b).

   (c) Is the following statement true or false: For waves travelling at the same speed, the longer the wavelength the greater the frequency? Why?

Model 2: Photons and Radiation Energy

Light can be thought of as an electromagnetic wave or electromagnetic radiation having a particular wavelength and frequency. In addition, Albert Einstein proposed almost a century ago that electromagnetic radiation can be viewed as a stream of particles known as photons, each of which has a particular amount of energy associated with it. Specifically, he proposed the following equation:

$E_{photon} = hf$ where h is called Planck's constant

Wavelengths, Frequencies, and Energies of Electromagnetic Radiation		
Wavelength (nm)	**Frequency (10^{14} s^{-1})**	**Energy (10^{-19} J)**
333.1	9.000	5.963
499.7	6.000	3.976
999.3	3.000	1.988

Regions of Electromagnetic Spectrum	
Region	**Wavelength Range**
Radiowave	3 km-30 cm
Microwave	30 cm-1 mm
Infrared (IR)	1 mm-800 nm
Visible	800 nm-400 nm
Ultraviolet (UV)	400 nm-10 nm
X-ray	10 nm-0.1 nm
Gamma ray	< 0.1 nm

The joule (J) is a unit of energy (1 J = 1 kg m^2/s^2)

1 nm = 10^{-9} m

4) A certain proton has a wavelength of 100 nm. In what region of the electromagnetic spectrum should this photon be classified (see right side of the table above)?

5) According to the data on the left side of the table above and the equation proposed by Einstein, what is the value of Planck's constant (including units)?

6) Based on the data on the left side of the table above and the relationship between speed, frequency, and wavelength, what is the speed of electromagnetic radiation (light waves)?

7) Quantities a and b are proportional when a = kb, where k is a constant. The two quantities are said to be inversely proportional when a = k/b.
   (a) Write the mathematical equation that relates the energy of a photon and its frequency. Is the energy of a photon proportional or inversely proportional to *f*?

   (b) Write the mathematical equation that relates the energy of a photon and its wavelength. Is the energy of a photon proportional or inversely proportional to λ?

8) Complete the table:

Energy (J)	**Wavelength (m)**	**Frequency (s^{-1})**	**Region of Spectrum**
		1.5×10^{14}	
	0.500×10^{-6}		
9.94×10^{-19}			
	1.00×10^{-9}		

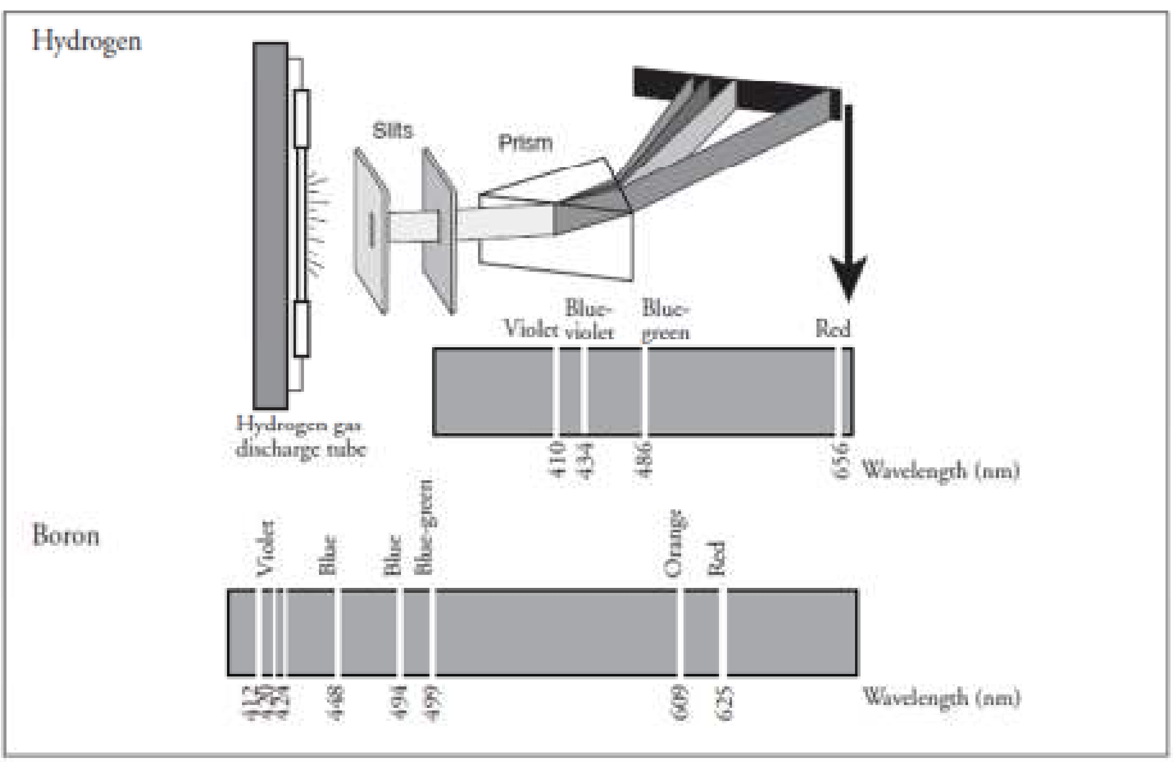
Atomic Spectra

From fireworks to stars, the color of light is useful in finding out what's in matter. The emission of light by hydrogen and other atoms has played a key role in understanding the electronic structure of atoms. Trace materials, such as evidence from a crime scene, lead in paint or mercury in drinking water, can be identified by heating or burning the materials and examining the color(s) of light given off in the form of bright-line spectra.

In Bohr's atomic model an atom's electrons are assigned to specific energy levels. The atom is in its <u>ground state</u> when the electrons occupy the lowest possible energy levels. When an electron absorbs sufficient energy it moves to a higher energy level to produce an <u>excited state</u>. When the electron releases the energy, it drops back to a lower energy level. The energy is released in the form of light. The wavelength of the emitted light indicates the difference in the energy of these two levels. Each wavelength of light corresponds to a specific color of light (which may or may not be visible). Consequently, atoms emit a characteristic set of discrete wavelengths - not a continuous spectrum.

Since each element has its own unique electron arrangement, the light that is emitted by the atoms produces an emission spectrum that can be used to identify the element. In other words, an atomic spectrum can be used as a "fingerprint" for an element because it is unique for each element and reflects the energy levels occupied by the electrons in an atom of the element.

Model 1: Emission Spectra for Hydrogen and Boron Atoms



1) Use color pencils to color the hydrogen and boron spectral lines with their respective spectra in Model 1.
2) The spectral lines for boron were produced using the same method as hydrogen. List three of the colors and corresponding wavelengths for boron's spectral lines as its light passes through a prism.

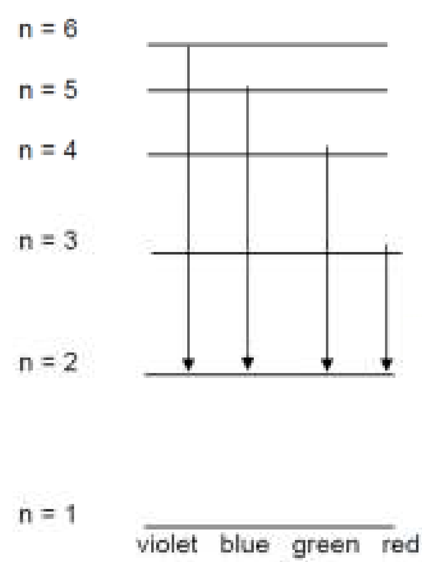
3) In the hydrogen spectrum in Model 1, which color of light corresponds to the shortest wavelength and to the longest wavelength?

4) In the hydrogen spectrum in Model 1, which color of light has the least energy and the most energy?

5) "The spectral lines for atoms are like fingerprints for humans." How do the spectral lines for hydrogen and boron support this statement?

Model 2: Energy Level Diagram for Hydrogen Atom

Niels Bohr modified Rutherford's Nuclear Atom model to explain how light interacted with the electrons in an atom to produce spectral lines. His model included electrons orbiting the nucleus at specific energy levels. Electrons absorb energy from various sources (electricity) when they move from lower energy levels (ground state) to higher energy levels (excited states). Energy is released as electrons return to their lower energy levels. If atoms only emit discrete wavelengths of light, then an atom's energy levels can only have discrete energies. The energy level diagram on the next page illustrates some of the energy levels found in a hydrogen atom, with arrows showing the corresponding electron transitions that produce its visible emission spectrum. The transitions shown are from excited states to the second energy level. Transitions to the first energy level (the ground state for the hydrogen atom) do not fall in the visible range (but in VU & IR range).

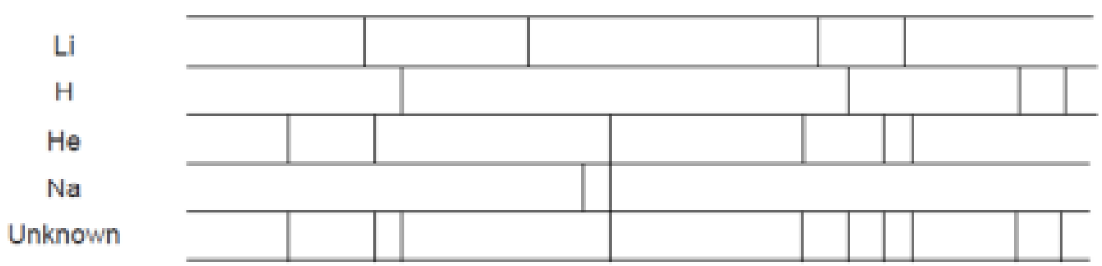


6) For the element hydrogen, which color (wavelength) of light is produced by the largest energy drop of an electron? Which color of light does the smallest energy drop of an electron produce? Explain your answer.

7) How does your answer compare to your answer in question 4?

8) Why do different elements emit different colors of light?

9) Below are the bright line spectra of four elements and the spectrum of an unknown gas from an experiment. What is the identity of the unknown gas or gases?



10) In a forensic investigation, it was suspected that a toxic chemical, which contains the element copper and accumulates in large amounts in hair as well as other tissues, poisoned the victim. Using your knowledge of emission spectra, design a method by which you can confirm the presence of this toxin in the victim.